Chemistry 129.01 Quiz #10 - 25 points

Name: Key

Show all your work!

1. (4 pts.) A buffer contains significant amounts of ammonia (NH_3) and ammonium nitrate (NH_4NO_3). (a) Write chemical equation showing how this buffer neutralizes added HCl. (b) Do you expect the pH of the buffer to decrease or increase after the addition of HCl? Why?

(a) $NH_{3\,(aq)} + HCl_{(aq)} \longrightarrow NH_4Cl_{(aq)}$

(b) The pH of the buffer will decrease after the addition of HCl. Ammonia neutralizes HCl and its concentration decreases while the concentration of NH_4^+ (weak acid of buffer) increases so pH is lower than initial pH.

2. (6 pts.) Two 40.0mL samples of unknown acids **A** and **B** are studied by titration with a 0.100 M NaOH solution. One sample is **aspirin** (acetylsalicylic acid, pK_a = 3.52), and the other is **vinegar** (acetic acid, pK_a = 4.74).
(a) Which titration curve corresponds to which acid? **Briefly** explain.
(b) What are the concentrations of the two acids samples?

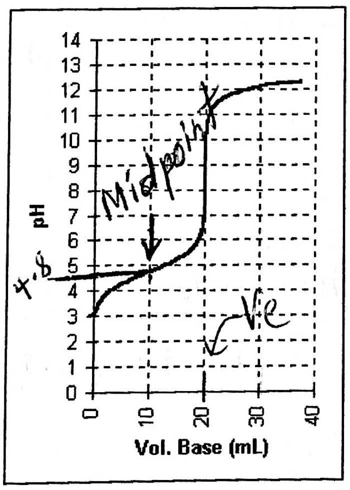
(b) $C_aV_a = C_bV_b$

$C_a = \dfrac{C_bV_b}{V_a}$

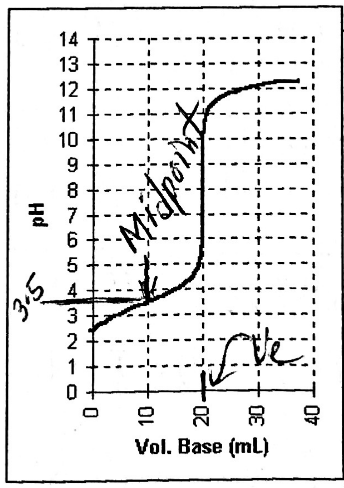
$C_a = \dfrac{(0.100M)(20.0mL)}{40.0mL}$

$= \underline{0.0500M}$

A



B



* Both acids' concentrations are the same. (Same volume of acid titrated & same equivalence volume).

(a) Equivalence Point (Ve) = 20.0 mL $\qquad \frac{1}{2}$Ve = 10.0mL (Midpoint)

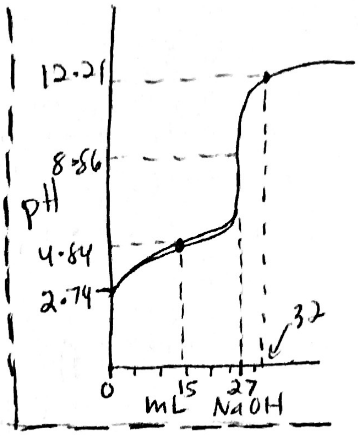
At $\frac{1}{2}$Ve, pH = pKa

A: pKa ≅ 4.8
Acetic Acid
(Vinegar)

B: pKa ≅ 3.5
Acetylsalicylic Acid
(Aspirin)

3. (15 pts.) Consider the titration of a 30.0mL sample of 0.180M acetic acid ($HC_2H_3O_2$) with a 0.200M ~~HCl~~ NaOH solution. Calculate the pH after the following volumes of the NaOH solution have been added: **0 mL, 15.00 mL, at the equivalence point and 32.00 mL.** Also, **sketch** a rough titration curve for this titration. Acetic Acid's K_a is 1.8×10^{-5}.

$$V_e = V_b = \frac{C_a V_a}{C_a} = \frac{(0.180\,M)(30.0\,mL)}{0.200\,M} = 27.0\,mL\ NaOH$$



$V_b = 0$ mL (Weak Acid Soln)

$$[H_3O^+] = \sqrt{K_a \times [HC_2H_3O_2]} = \sqrt{(1.8\times10^{-5})(0.180)} = 1.8\times10^{-3}M$$

$$pH = -\log(1.8\times10^{-3}) = \underline{\underline{2.74}}$$

$V_b = 15.00$ mL (Buffer Region)

	$HC_2H_3O_2$ +	NaOH →	$NaC_2H_3O_2$	+ H_2O
BA	5.40 mmol			
A		3.00 mmol		
AA	2.40 mmol	0	3.00 mmol	

$$pH = pK_a + \log\frac{[C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$pH = -\log(1.8\times10^{-5}) + \log\frac{(3.00)}{(2.40)}$$

$$pH = \underline{\underline{4.84}}$$

$V_b = 27.0$ mL (Equivalence Point – Weak base Soln)

	$HC_2H_3O_2$ +	NaOH →	$NaC_2H_3O_2$	+ H_2O
BA	5.40 mmol			
A		5.40 mmol		
AA	0	0	5.40 mmol	

$$[C_2H_3O_2^-] = \frac{5.40\,mmol}{57.0\,mL} = 0.0947\,M$$

$$K_b = \frac{1.0\times10^{-14}}{1.8\times10^{-5}} = 5.6\times10^{-10}$$

$$[OH^-] = \sqrt{K_b \times [C_2H_3O_2^-]} = \sqrt{(5.6\times10^{-10})(0.0947)} = 7.3\times10^{-6}M$$

$$pOH = -\log(7.3\times10^{-6}) = 5.14 \qquad pH = 14.00 - 5.14 = \underline{\underline{8.86}}$$

$V_b = 32.00$ mL (Excess Strong Base)

	$HC_2H_3O_2$ +	NaOH →	$NaC_2H_3O_2$	+ H_2O
BA	5.40 mmol			
A		6.40 mmol		
AA	0	1.00 mmol	5.40 mmol	

$$[OH^-] = [NaOH] = \frac{1.00\,mmol}{62.0\,mL} = 0.0161\,M$$

$$pOH = -\log(0.0161) = 1.792$$

$$pH = 14.00 - 1.792 = \underline{\underline{12.21}}$$